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ONTARIO

DEPARTMENT OF EDUCATION

CHEMISTRY

Senior Division
Grade 12
Five-Year Program
All Branches

Replacing the course in Chemistry for Grade 12
as printed in Curriculum S.17, Science and Agricultural Science

For Introduction in September, 1966

ISSUED BY AUTHORITY OF
THE MINISTER OF EDUCATION

GRADE 12 CHEMISTRY

This course of study forms the first part of a two-year program in chemistry for Ontario schools. It is a preparation for more advanced study in Grade 13, but it is also designed to be an adequate introduction to chemistry for those students who do not undertake further study in the subject.

An important feature of the course is that great emphasis is placed on attempts to understand the nature of elements and compounds and the reactions they undergo. Theories of the structure of atoms and aggregates of atoms are introduced early in the course so that these ideas may be used in subsequent discussions. Wherever possible each topic is illustrated by experiment and discussed in terms of theoretical principles. It is intended that experiment and discussion develop together throughout the course. To facilitate this it is essential that one double period be available each week for the longer experiments. The course is based on six periods each week, as suggested in Circular H.S.I., "Requirements for Diplomas and Statements of Standing".

It cannot be too strongly emphasized that the aim of the course is to create an understanding of the subject, rather than to provide information to be memorized. The course provides an introduction to the experimental study of chemistry, to the simple theoretical concepts of structure, bonding, reactions, and solutions, and to the language of the subject. To achieve understanding, the student must learn to think and to experiment for himself. To communicate his ideas to others he must become familiar with the language of the subject and be able to illustrate theory by the use of facts.

None of the material of this course will be repeated as such in Grade 13, but the ideas will be used as a basis of further study. The background provided in this course will make possible in Grade 13 a meaningful introduction to the study of more advanced atomic structure and the periodic classification, to bonding, structure and chemical energy, and to topics such as kinetics, equilibrium, and oxidation-reduction.

UNIT HEADINGS

- Unit I INTRODUCTION
- Unit II THE STRUCTURE OF INDIVIDUAL ATOMS
1. Fundamental Particles
 2. The Atom
 3. Ion Structure
- Unit III THE STRUCTURE OF AGGREGATES OF ATOMS
1. Occurrence
 2. Categories
 3. Ionic Crystals
 4. Molecules
 5. Covalent Crystals
 6. Metals
- Unit IV STATES OF MATTER
1. Some Properties of Solids, Liquids and Gases
 2. Solids
 3. Liquids
 4. Gases and the Kinetic-Molecular Theory
- Unit V OXYGEN AND HYDROGEN: AN INTRODUCTION TO CHEMICAL REACTIONS
1. Nature of Chemical Reactions
 2. Chemical Equations
 3. Sources of Oxygen
 4. Some Reactions of Oxygen
 5. Some Properties of Oxides
 6. Importance of Oxygen
 7. Sources of Hydrogen
 8. Some Reactions of Hydrogen
 9. Relative Activities of Metals
- Unit VI THE MOLE, ATOMIC WEIGHTS, AND MOLECULAR WEIGHTS
1. The Mole
 2. Atomic Weights and Gram-Atomic Weights

3. Molecular Weights and Gram-Molecular Weights
4. Problems

Unit VII FORMULAS, NOMENCLATURE, AND EQUATIONS

1. Formulas and Nomenclature
2. Valence
3. Equations

Unit VIII WATER AND SOLUTIONS

1. Water
2. Solutions

Unit IX IONS IN AQUEOUS SOLUTION

1. Occurrence of Ions
2. Electrolysis
3. Some Redox Reactions
4. Properties of Solutions Containing Ions in Common
5. Neutralization
6. Some Factors Affecting the Extent of Chemical Reactions

Unit X ELEMENTS OF GROUP 2

1. Magnesium and Calcium
2. Beryllium, Strontium, Barium, and Radium
3. Gradations in Reactivities

Unit XI ELEMENTS OF GROUP 7

1. The Halides
2. The Halogens
3. Similarities and Gradations in Properties

Unit XII PERIODIC CLASSIFICATION OF ELEMENTS

General Features of the Periodic Table

Unit I: Introduction (1½ weeks)

1. Scope of chemistry and its relationships with other sciences and mathematics
Nature of the scientific method
2. The chemist's view of matter as made up of a finite number of pure substances, sometimes occurring isolated, but more often occurring in mixtures
3. Properties as characteristics which identify pure substances, e.g., colour, density, boiling point, melting point, conduction of electricity, etc.
4. Classification of pure substances as elements (characteristics, including reference to limit of decomposition) or compounds (characteristics, including constant composition)
5. Physical change and chemical change
Synthesis and decomposition

Experiments

- (a) Boiling point of water, or carbon tetrachloride, or methanol
- (b) Melting point of naphthalene, or paradichlorobenzene
- (c) Quantitative decomposition of various samples of mercury(II) oxide
- (d) Synthesis of a sulphide (copper and zinc)

Unit II: The Structure of Individual Atoms (1½ weeks)

Note: Students should be trained in the recognition and use of symbols of the elements e.g., Na, and of individual nuclides, e.g., ${}_{17}\text{Cl}^{35}$; Cl-35

1. Fundamental Particles: electron, proton, neutron
 - (a) Absolute mass
 - (b) Approximate scale of atomic masses, based on proton or neutron as unity
 - (c) Charge
 - (d) Approximate size

2. The Atom

- (a) Nuclear atom (Rutherford)
- (b) Atomic number and mass number
- (c) Electron arrangements, e.g., 2,8,8,2 (Note: Consider elements only among the first twenty)
- (d) Schematic diagrams of atoms of atomic number 1 to 20 from given atomic numbers and mass numbers
- (e) Isotopes
 - (i) Occurrence of different forms of the same element, differing in mass number
 - (ii) Constancy or near constancy of isotopic composition in nature
 - (iii) Atomic weight of an element from the mass and relative abundance of its isotopic forms
 - (iv) Precise definition of the atomic-weight scale in terms of carbon-12

3. Ion Structure

- (a) Relative stability of electron structures of the inert gases, and the tendency of other atoms to gain or lose electrons to form ions with such structures
- (b) Positive ions formed by removal of electrons from neutral atoms
Negative ions formed by addition of electrons to neutral atoms
- (c) Tendency of atoms of metallic elements to be electropositive, i.e., to lose electrons to become positive ions
Tendency of atoms of non-metallic elements to be electronegative, i.e., to gain electrons to become negative ions

Note: Ions, whether positive or negative, are formed readily only by the more reactive elements

Unit III: The Structure of Aggregates of Atoms (1½ weeks)

- 1. Isolated atoms do not usually occur in nature (exceptions being the inert gases: helium, neon, etc.)
- 2. Aggregates of atoms in four categories: ionic crystals, molecules, covalent crystals, metals

3. Ionic Crystals

- (a) Solids consisting of regular arrays of ions, neutrality resulting from the relationship between the charges on the ions and the ratio of positive to negative ions, e.g.,
 NaCl (Na^+Cl^-); MgO ($\text{Mg}^{2+}\text{O}^{2-}$); Na_2O ($2\text{Na}^+\text{O}^{2-}$), CaF_2 ($\text{Ca}^{2+}2\text{F}^-$)
- (b) The charge on the ion as the electrovalence (ionic valence) of the element, e.g., the calcium in CaF_2 has an electrovalence of 2+ and the fluorine an electrovalence of 1—

Note: In indicating the electrovalence, the number before the sign is preferred, e.g., Al^{3+} rather than Al^{+3} or Al^{+++} ; when the electrovalence is 1, either positive or negative, the number is usually omitted, e.g., Na^+ rather than Na^{1+}

- (c) The force of attraction of unlike charges binding the solid together

4. Molecules

- (a) Combinations of atoms as molecules, represented by combinations of symbols called formulas, e.g., H_2 , HF , NH_3 , CHCl_3
- (b) Covalent bond: two nuclei held together by a pair of shared electrons (single bond) giving a diatomic molecule

Prototype as the molecule of hydrogen, H_2 ; other examples: F_2 , Cl_2 , HF , HCl

- (c) Chemical bond between atoms represented by a pair of dots or a single dash, e.g., $\text{H}:\text{Cl}$ or $\text{H}-\text{Cl}$
- (d) The possibility of the atom forming more than one covalent bond if it has, or can have, more than one unpaired electron, e.g., H_2O , NH_3 , CH_4

Occurrence of double or triple bonds if more than one pair of electrons join the two atoms, e.g., $\text{H}_2\text{C}=\text{CH}_2$, $\text{O}=\text{C}=\text{O}$, $\text{HC}\equiv\text{CH}$, $\text{N}\equiv\text{N}$

- (e) Covalence: the number of bonds formed by one atom, e.g., the nitrogen in ammonia having a covalence of three, and the carbon in carbon dioxide having a covalence of four

- (f) The bond representing the force holding the atoms together, against which work must be done to break the molecule into its constituent atoms

Note: The distinction between ionic and covalent bonding is not always clear

5. Covalent Crystals

- (a) Composed of atoms linked by covalent bonds, but of such vast numbers of atoms that the term "molecule" is inappropriate
- (b) Occurrence as solids, and the assemblages of atoms, covalently linked, comprising the single piece of that solid, e.g., diamond and silica
- (c) The fact that the formula of such a substance does not describe a molecule, but merely gives the ratio in which the atoms are present, e.g., SiO_2

6. Metals

- (a) The properties of metals, such as electrical and thermal conductivity, reflection of light, and others, leading to the picture of the metallic state as an assemblage of partially ionized atoms with the liberated electrons binding the atoms together
- (b) The freedom of these electrons to move among the stationary atoms, and carry an electric current through the metal when an electric potential is applied

7. Use of models to illustrate the structure of various aggregates, e.g., sodium chloride, methane, ammonia, water, hydrogen, diamond, copper

Unit IV: States of Matter (3½ weeks)

1. Some Properties of Solids, Liquids and Gases

- (a) Particulate nature
- (b) Forces of attraction between particles
- (c) Types of motion: translational, vibrational, rotational
- (d) Elasticity of collisions
- (e) Distance apart of particles

2. Solids

- (a) Both shape and volume definite at a given temperature

Ordered pattern (crystal lattice) of particles
Motion of particles vibrational only

- (b) Types of structural units: similar atoms, e.g., diamond, copper; different atoms, e.g., silicon dioxide; molecules, e.g., ice, sulphur; ions, e.g., sodium chloride

Note: Distinctions between types of structural units are not always as sharp as indicated by the examples

- (c) Vibrational Energy and Temperature

(i) Change of volume with change in temperature

(ii) Temperature indicative of average kinetic energy of particles

(iii) Melting point is the temperature at which the particles have sufficient vibrational energy to break loose from their ordered structure

Freezing and freezing point

(iv) Heat of fusion as the energy required to break the particles loose

Dependence of the quantity of energy on the number of particles involved and the magnitude of the forces that hold the particles in the solid state

(v) Sublimation (brief reference)

- (d) Reference to non-crystalline solids, such as glass and plastics, and to their lack of ordered structure, of sharp melting point, of plane surfaces, and of definite interfacial angles

Observation of examples of surfaces of fractured crystalline and non-crystalline solids

3. Liquids

- (a) Shape indefinite but volume definite at a given temperature

Translational motion of particles (atoms, molecules, ions) relative to each other but still essentially in contact

Similarity of volume as solid or liquid (similar density)

(b) Energy and Temperature

- (i) Change of volume with change in temperature
- (ii) Temperature indicative of average kinetic energy of particles
- (iii) Distribution of speeds of particles and hence distribution of kinetic energies of particles

(c) Vaporization

- (i) Evaporation as escape from the liquid of those particles with sufficient energy
Cooling effect due to decrease in average kinetic energy of particles which remain in liquid state
- (ii) Heat of vaporization as energy required to give particles sufficient energy to escape from the liquid
Dependence of the quantity of energy on the number of particles vaporizing and the magnitude of the cohesive forces between the particles
Condensation and the release of heat

(d) Vapour Pressure

- (i) Pressure exerted by escaped particles when in equilibrium with the liquid
- (ii) Relationship between vapour pressure and temperature
- (iii) Relationship between vapour pressure and boiling point
High boiling points of metals, e.g., Fe, Cu, and low boiling points of molecular substances, e.g., H_2O , CCl_4 , related to forces between particles

4. Gases and the Kinetic-Molecular Theory

Note: The kinetic-molecular theory, in which both the forces between molecules of a gas and the volume of the molecules themselves are ignored, is strictly valid only for a so-called ideal or perfect gas. Real gases become more ideal the higher

their temperature and the lower their pressure. In this Grade 12 introductory treatment, gases will be treated as if they were ideal.

- (a) Both shape and volume indefinite and determined by the container

High compressibility, low density, and rapid diffusion

Translational motion of particles

Pressure resulting from collision of particles with walls of container, and dependent on frequency of impact per unit of area and on average kinetic energy of particles

- (b) Relationship of speed, kinetic energy, and temperature (Kelvin) ,

Note: At the same temperature, particles of all gases have the same average kinetic energy.

- (c) Avogadro's Principle

- (d) Boyle's Law

- (i) Experiment to illustrate the law

- (ii) Explanation of the law using the kinetic-molecular theory (average kinetic energy of particles constant at constant temperature; relationship between pressure, frequency of collisions, concentration of particles, and volume)

- (e) Charles' Law

- (i) Experiment to illustrate the law
Graphical interpretation of results

- (ii) Explanation of the law using the kinetic-molecular theory as in 4 (d) (ii)

- (f) Problems involving

- (i) Volume, density, and pressure at constant temperature

- (ii) Volume, density, and temperature at constant pressure

- (iii) Volume, and density with simultaneous changes in temperature and pressure

Experiments

- (a) Demonstration of sublimation
- (b) Demonstration of vapour pressure of water and of benzene, and the effect of temperature on vapour pressure
- (c) Demonstration of relative rates of diffusion of different gases, e.g., hydrogen chloride and ammonia
- (d) Boyle's Law
- (e) Charles' Law

Unit V: Oxygen and Hydrogen:

Introduction to Chemical Reactions (4½ weeks)

1. Chemical reaction as a rearrangement of atoms involving the formation of a new product or products, and the release or absorption of energy (commonly as heat)
Meaning of exothermic and endothermic
2. The chemical equation as a concise statement of a chemical reaction, showing conservation of atoms as the basis for the balanced equation
3. Sources of Oxygen
 - (a) Natural occurrence: in water and as complex compounds in rocks and living things
The role of photosynthesis in providing elemental oxygen to the air
 - (b) Laboratory preparation: heating oxides and chlorates, e.g., mercury (II) oxide and potassium chlorate
 - (c) Commercial production: distillation of liquid air and electrolysis of water
4. Some Reactions of Oxygen (Oxidations)
 - (a) Formation of oxides by combustion of elements, e.g., sodium, magnesium, iron, carbon, phosphorus, sulphur
 - (b) Combustion of carbonaceous substances, e.g., natural gas or propane, alcohol, paraffin, to form carbon dioxide and water

- (c) Energy release associated with combustion, attention being drawn to the necessity of providing energy to initiate some reactions

Dependence of the quantity of energy released on the particular reaction and on the amount of the substance burned

- (d) Low-temperature reactions, e.g., oxidation in air of iron, sodium, phosphorus

Spontaneous combustions initiated by oxidation at low temperature, e.g., in oily rags, coal piles

Formation and value of protective metal oxide coatings, e.g., aluminum oxide

5. Some Properties of Oxides

- (a) Differences in properties of aqueous solutions of oxides of metals and of non-metals
- (b) Introduction to the concepts: acid, base

6. Importance of Oxygen

- (a) In respiration and the production of energy in the living cell
Functions of this energy
- (b) In combustions, in household and industrial furnaces, to produce heat
- (c) Medical uses of oxygen-enriched air, e.g., oxygen tent, oxygen inhalator
- (d) Pure oxygen for increasing a rate of combustion, e.g., in rockets, steelmaking, oxy-acetylene torch

7. Sources of Hydrogen

- (a) Natural occurrence as plentiful in the sun and other stars, but rare on the earth except in the form of compounds
- (b) Laboratory preparation by the action of water or steam on metals and by the action of acids on metals
- (c) Commercial production: cracking of petroleum or natural gas, reaction of hot coke with steam, electrolysis of water

8. Some Reactions of Hydrogen

- (a) With oxygen (explosive properties)
- (b) With oxides of metals, e.g., CuO , Fe_3O_4

Reversibility of the reaction involving iron oxide and hydrogen

9. Relative activities of metals as shown by the vigour of their reactions with oxygen and by their ability to displace hydrogen from acids

Experiments

- (a) Preparation of oxygen by heating potassium chlorate
- (b) Catalytic effect of manganese dioxide on the decomposition of potassium chlorate
- (c) Combustion of sodium, magnesium, iron, carbon, phosphorus, and sulphur in oxygen; the nature of solutions of the products formed
- (d) Detection of the products of combustion of carbonaceous substances
- (e) Determination of heat of combustion of candle material
- (f) Demonstration of the spontaneous combustion of white phosphorus
- (g) Demonstration of the protective oxide coating on aluminum and its removal using mercury(II) chloride
- (h) Reaction of water with potassium, sodium, and calcium and of steam with magnesium, iron, and copper
- (i) Reaction of hydrochloric acid and of dilute sulphuric acid with magnesium, iron, and copper
- (j) Demonstration of the combustion of hydrogen
- (k) Reaction of copper(II) oxide with hydrogen
- (l) Demonstration of the reaction of hydrogen with magnetic iron oxide

Unit VI: The Mole, Atom Weights, and Molecular Weights (3½ weeks)

1. The Mole: 6.02×10^{23} particles (atoms, molecules, ions)
Recognition of 6.02×10^{23} as Avogadro's Number, designated N
2. Atomic Weights and Gram-Atomic Weights
 - (a) The method of determination of mass numbers and atomic weights by the mass-spectrographic method
 - (b) Gram-atomic weight: weight in grams of one mole of atoms (N atoms) of the element
3. Molecular Weights and Gram-Molecular Weights
 - (a) Distinction between molecular weight and gram-molecular weight
 - (b) Gram-molecular weight: weight, in grams, of one mole of molecules (N molecules) of the compound
 - (c) Avogadro's Principle
 - (d) Gram-molecular volume (molar volume)
 - (e) Determination of gram-molecular weights
 - (i) from formula of compound and gram-atomic weights
 - (ii) from mass, volume, temperature, and pressure of a gas
4. Problems Involving Calculation
 - (a) Composition by weight from formula
 - (b) Simplest formula from composition by weight
 - (c) Molecular formula from simplest formula and molecular weight
 - (d) Number of atoms of an element from a given weight of a compound of given formula
 - (e) Gram-molecular weight of a gas from its mass, volume, temperature, and pressure
 - (f) One of mass, volume, temperature, and pressure of a gas from the other three of these and the gram-molecular weight
 - (g) Concentration of molecules in a given weight and volume of a gas

Experiments

- (a) Determination of the volume occupied by one gram-molecular weight of oxygen at STP
- (b) Determination of the ratio by weight of equal volumes of different gases

Unit VII: Formulas, Nomenclature, and Equations (3½ weeks)

1. Formulas and Nomenclature

- (a) Names and symbols of common elements
Order of symbols in formula of a binary compound, e.g., NaCl, not ClNa
Use of suffix “ide” for binary compounds, e.g., oxide, hydride, chloride
- (b) Existence of compound ions, e.g., SO_4^{2-} , NO_3^- , and names of common ones
- (c) Nomenclature of simple salts, e.g., “-ic” acids yield “-ate” salts, etc.

2. Valence

- (a) Valence of common elements and compound ions
- (b) Valence of elements in binary compounds from formulas including elements with more than one valence
Preference for IUPAC nomenclature, e.g., iron(II) preferred to ferrous
- (c) Writing names from formulas, and formulas from names
- (d) Valence and atomic structure, i.e., relationship between atomic structure and chemical combination

3. Equations

- (a) Balancing equations
- (b) Information derivable from balanced equations
- (c) Problems on equations involving relationships of weight-weight, weight-volume, volume-volume

Experiments

- (a) Determination of the formula of
—mercury(II) oxide (data from Unit I—
Experiment (c))
or
—copper(I) sulphide by heating copper wire
with an excess of sulphur
- (b) Determination of the number of moles of silver
produced from a solution of silver nitrate by one
mole of copper (verification of chemical equation)
- (c) Determination of the number of moles of hydrogen
produced by one mole of magnesium (verification
of chemical equation)
- (d) Demonstration of the ratio by volume in which hy-
drogen and oxygen react

Unit VIII: Water and Solutions (1½ weeks)

1. Water

- (a) Occurrence and importance
- (b) Physical properties, including density, vapour
pressure, freezing point, boiling point, and the
effect of temperature on density and on vapour
pressure
- (c) Use as a basis of standards: litre, calorie, specific
gravity, thermometer scales
- (d) Composition by volume (as revealed by electro-
lysis), and composition by weight (as revealed by
reduction of an oxide by hydrogen)
- (e) Hydrates
 - (i) Qualitative detection of water in copper(II)
sulphate pentahydrate
 - (ii) Quantitative determination of water in bar-
ium chloride dihydrate
 - (iii) Efflorescence and efflorescent hydrates
 - (iv) Problems involving formulas of hydrates
- (f) Hygroscopic and deliquescent substances

2. Solutions

- (a) Characteristics including homogeneity, permanence, variability of composition, and limits of miscibility
- (b) Types of solutions
Solids, liquids, and gases as solutes and as solvents
- (c) Methods of expressing concentration and related problems, specifically: percentage by weight, weight of solute per 100 g. of solvent, weight of solute per litre of solution, and molarity

Experiments

- (a) Decomposition of copper(II) sulphate pentahydrate
- (b) Quantitative determination of water in barium chloride dihydrate (verification of formula)

Unit IX: Ions in Aqueous Solution (4 weeks)

1. Occurrence of Ions

- (a) Comparison of conduction of electricity by metals and by aqueous solutions
- (b) Definition of electrolyte and non-electrolyte
- (c) Sources of mobile ions
 - (i) The dissociation of ions of electrovalent compounds (in solution or fused)
 - (ii) The reaction of a covalent compound with water to produce ions (ionization)

2. Electrolysis

- (a) Definition of electrode, anode, cathode, and electrolysis
- (b) Electrolysis of hydrochloric acid, copper sulphate solution, sodium chloride solution, and water containing some potassium nitrate or sulphuric acid
- (c) Electrolysis as a phenomenon involving electron transfer
Identification of electrode processes (electron transfer between ions and electrodes) with oxidation and reduction.

Reduction at the cathode and oxidation at the anode

Electrode reaction not necessarily involving the ions from the solute

Electrochemical series

3. A review of oxidation and reduction reactions considered previously, e.g., metals with acids or with oxygen, in which oxidation and reduction are considered as electron transfer processes (without electrodes)

Redox as the general name for these processes taken together

Ionic equations for such reactions, e.g., $\text{Mg} + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2$

Relevance of the electrochemical series

4. Properties of Solutions Containing Ions in Common

(a) Salts of coloured ions

(b) An acid as a substance that produces hydrogen ions (or hydronium ions)

(c) A base as a substance that produces hydroxide ions

5. Neutralization

(a) The reaction of hydroxide ions and hydrogen ions (hydronium ions) forming water, the ions of a salt remaining in solution

(b) Reference to the step-wise neutralization of polyprotic acids

(c) Review of nomenclature (Unit VII—1)

IUPAC nomenclature of acid salts, e.g., sodium hydrogen carbonate preferred to sodium bicarbonate

6. Some Factors Affecting the Extent of Chemical Reactions

(a) Formation of an “insoluble” (very slightly soluble) salt

Use of solubility rules for common salts in anticipating such precipitation

(b) Formation of gases of low solubility, e.g., H_2

(c) Formation of a slightly ionized substance, e.g., H_2O

- (d) Relative position of metals in the electrochemical series, e.g., $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ occurring but $\text{Cu} + \text{Zn}^{2+} \rightarrow \text{Cu}^{2+} + \text{Zn}$ not occurring

Experiments

- (a) Conductivity of aqueous solutions of hydrogen chloride, sodium chloride, sodium hydroxide, methanol, sucrose, and a solution of hydrogen chloride in carbon tetrachloride
- (b) Electrolysis as indicated in 2(b)
- (c) Illustration of 4(a) by a demonstration of colour of dilute solutions of several salts of each of copper, cobalt, and nickel
- (d) Properties common to (i) acids, (ii) bases
- (e) Neutralization
- (f) Formation of the following substances to illustrate
 - (i) 6(a) PbI_2 , BaSO_4 , CuS
 - (ii) 6(b) H_2 , CO_2 , H_2S
- (g) Relative position of zinc, lead, copper, and silver in the electrochemical series as shown by the displacement reactions in suitable solutions
- (h) Determination of the relative position of magnesium, zinc, iron, and copper in the electrochemical series by the reaction of these metals with dilute hydrochloric acid

Unit X: Elements of Group 2 (2 weeks) (The Alkaline Earths)

1. Magnesium and Calcium

- (a) Similarities of electron arrangements and valence (2,8,2; 2,8,8,2)
- (b) Occurrence
- (c) Production by electrolysis
- (d) Physical properties of the elements
- (e) Reactions with oxygen, water, and acids
- (f) Oxides
 - (i) Formation; direct union, decomposition of carbonates

- (ii) Structure
 - (iii) Reaction with water
 - (iv) Uses and related properties
 - (g) Carbonates
 - (i) Formation from oxides, hydroxides, hydrogen carbonates
 - (ii) Reaction with acids
 - (iii) Uses and related properties
 - (h) Hardness of water
 - (i) Temporary and permanent
 - (ii) Softening by precipitation and by ion exchange
 - (iii) Comparison of reaction with soaps and detergents
 - (i) Sulphates: the preparation and hydration of plaster of Paris
2. Brief reference to beryllium, strontium, barium, and radium to draw attention to the similarity of their electron structure (outer shell)
 3. Gradations in reactivities of elements in group 2 as related to size of their atoms, electron arrangements, and ease of removal of valence electrons (ionization potential)

Notes Regarding Experiments

1. It is expected that Unit X will be treated experimentally
2. Experimentation should illustrate similarities and gradations of properties where they exist, e.g., reaction of metals with water and the decomposition of carbonates

Unit XI: Elements of Group 7 (2½ weeks)

1. The Halides
 - (a) Occurrence
 - (b) Study of the properties of the halides of sodium and potassium including colour, solubility, and crystal structure (review)

- (c) Study of the properties of the halides of silver including precipitation, effect of light, and solubility in
 - (i) water
 - (ii) aqueous ammonia
 - (d) The preparation of the halides of hydrogen and a study of their properties including aqueous solutions and their stability as illustrated by the action of concentrated sulphuric acid on halides of potassium
2. The Halogens
- (a) Production of the halogens by the oxidation of the halides
 - (b) Preparation of chlorine, bromine, and iodine
Their properties including state, colour, solubility in water and in carbon tetrachloride, bleaching action, and displacement
 - (c) Fluorine and the properties and uses of some of the important compounds of fluorine, e.g., sodium fluoride and hydrogen fluoride
 - (d) Astatine (brief reference)
3. Similarities and Gradations of Properties of Elements of Group 7 Including:
- (a) High ionization potential of halogens compared to metals of group 2
 - (b) Outer electron arrangements of ions and atoms
 - (c) Sizes of ions and atoms
 - (d) Valence
 - (e) Halogens as to state, colour, and oxidizing power related to electron arrangements
 - (f) Halides as to state, colour, solubility, and stability of the halides of hydrogen and silver

Notes Regarding Experiments

1. It is expected that Unit XI will be treated experimentally with the exception of the sections on fluorine and astatine
2. These experiments should bring out the differences in properties of the halide ion and the halogen molecule
3. Experiments should also be used to illustrate similarities and gradations of properties where they exist

Unit XII: Periodic Classification of the Elements

(1/2 week)

General Features of the Periodic Table

- (a) An arrangement of the elements based on atomic structure
- (b) Placement of elements with similar arrangements of external electrons into vertical rows called groups
- (c) Horizontal rows in this arrangement, called periods, corresponding to gradual filling of shells of electrons
- (d) Similarity in properties of elements in the same group with gradual changes associated with changes in atomic size
- (e) Changes in properties of elements in the same period resulting from increases in nuclear charge and in numbers of external electrons